IB Chemistry | Brakke | ECA – Topic 08

# T08D04 : pH and the Equilibrium Constant for Water

Name ........................................

**Part A: Essentials**

Water will ionize(dissociate) into its ions according to the following equation

$H_2O\,(l) \rightarrow H^+\,(aq) + OH^-\,(aq)$

The equilibrium constant for water is therefore

$K_w = [H^+][OH^-]$ and its value is =

$1 \times 10^{-14}$ at 25 degrees C. This value means that Whether there is an excess of H+, OH−, or equal amounts, the sum of the parts is still equal to $1\times10^{-14}$.

In comparison the equilibrium value for HCl is exceedingly large $H^+ > HCl$

**Part B: The Formulas and Problems**

The formula for pH is :

$-\log[H^+] = pH$ * where [H+] is the concentration of H+ ions in mol/L (molarity for strong acids)

The additional formulas derived from the pH and $K_w$ formulas are:

$pH + pOH = 14$

$-\log[OH^-] = pOH$

$10^{-pH} = [H^+]$ $10^{-pOH} = [OH^-]$

1. Using this formula calculate the pH of the following:

a. $[H^+] = 1.67 \times 10^{-4}$ 3.78

b. $[H^+] = 1.1 \times 10^{-5}$ 4.96 $-\log[H^+] = pH$

c. $[H^+] = 1.8 \times 10^{-11}$ 10.74

d. $[H^+] = 5.5 \times 10^{-6}$ 5.26

2. Calculate the **pH** of the solutions that have the following concentrations of <u>hydrogen</u> or <u>hydroxide</u> ions:

a. $[H^+] = 1.67 \times 10^{-4}$ 3.78

b. $[OH^-] = 1.1 \times 10^{-5}$ 9.04 $-\log[H^+] = pH$

$-\log[OH^-] = pOH$

c. $[H^+] = 1.8 \times 10^{-11}$ 10.74

d. $[OH^-] = 5.5 \times 10^{-6}$ 8.74 $14 - pOH = pH$

3. Calculate the hydrogen ion concentration, $[H^+]$, and the hydroxide ion concentration , [OH-] of each of the f ollowing solutions given their pH:

| | pH | pOH | $[H^+]$ | $[OH^-]$ |
|---|---|---|---|---|
| a. | 2.9 | 11.1 | $1.26\times10^{-3}$ | $7.94\times10^{-12}$ |
| b. | 5.5 | 8.5 | $3.16\times10^{-6}$ | $3.16\times10^{-9}$ |
| c. | 10.1 | 3.9 | $7.94\times10^{-11}$ | $1.26\times10^{-4}$ |
| d. | -1.4 | ~14 | $\sim1\times10^{0}$ | $\sim1\times10^{-14}$ |

4. Fill in the following table:

| $[H^+]$ | $[OH^-]$ | pH | pOH |
|---|---|---|---|
| 1.26E-2 | 7.94E-13 | 1.9 | 12.1 |
| 6.67E-11 | $1.5 \times 10^{-4}$ | 10.18 | 3.82 |
| 1E-14 | 1E0 | 14 | -0.60 |
| $2.3 \times 10^{-2}$ | 4.35E-13 | 1.64 | 12.36 |
| 3.16E-6 | 3.16E-9 | 5.5 | 8.5 |
| 0.000155 | 6.45E-11 | 3.81 | 10.19 |

5. Calculate the pH, pOH, $[H^+]$, and $[OH^-]$ concentration of each of the following:

a. 1.0 M solution of $HNO_3$ - acid

$-\log[1.0] = pH = 0$ $[H^+] = 1\times10^{0}$

$pOH = 14$ $[OH^-] = 1\times10^{-14}$

b. 0.65 M solution of NaOH - base

$-\log[0.65] = pOH = 0.19$ $[OH^-] = 6.5\times10^{-1}$

$pH = 13.81$ $[H^+] = 1.5\times10^{-14}$

c. 13.45 g of $H_2SO_4$ placed in water to make 325 mL of solution. [Diprotic acid but only the first ionization step should be used.] - acid

$H_2SO_4 \rightarrow HSO_4^- + H^+$

$13.45\,g\ H_2SO_4 \times \frac{1\,mol\ H_2SO_4}{98.09\,g\ H_2SO_4} = 0.137\,mol\ H_2SO_4$

$\frac{0.137\,mol}{0.325\,L} = \frac{mol}{L} = \text{Molarity}$

$= 0.422\,M$

$-\log[0.442] = 0.38 = pH$ $[H^+] = 0.442$

$13.62 = pOH$ $[OH^-] = 2.26\times10^{-14}$